ATAR Notes

Chemistry 1/2

Unit 1 Head Start

January Lecture Series

Presented by:

Josh Hamilton

Welcome!

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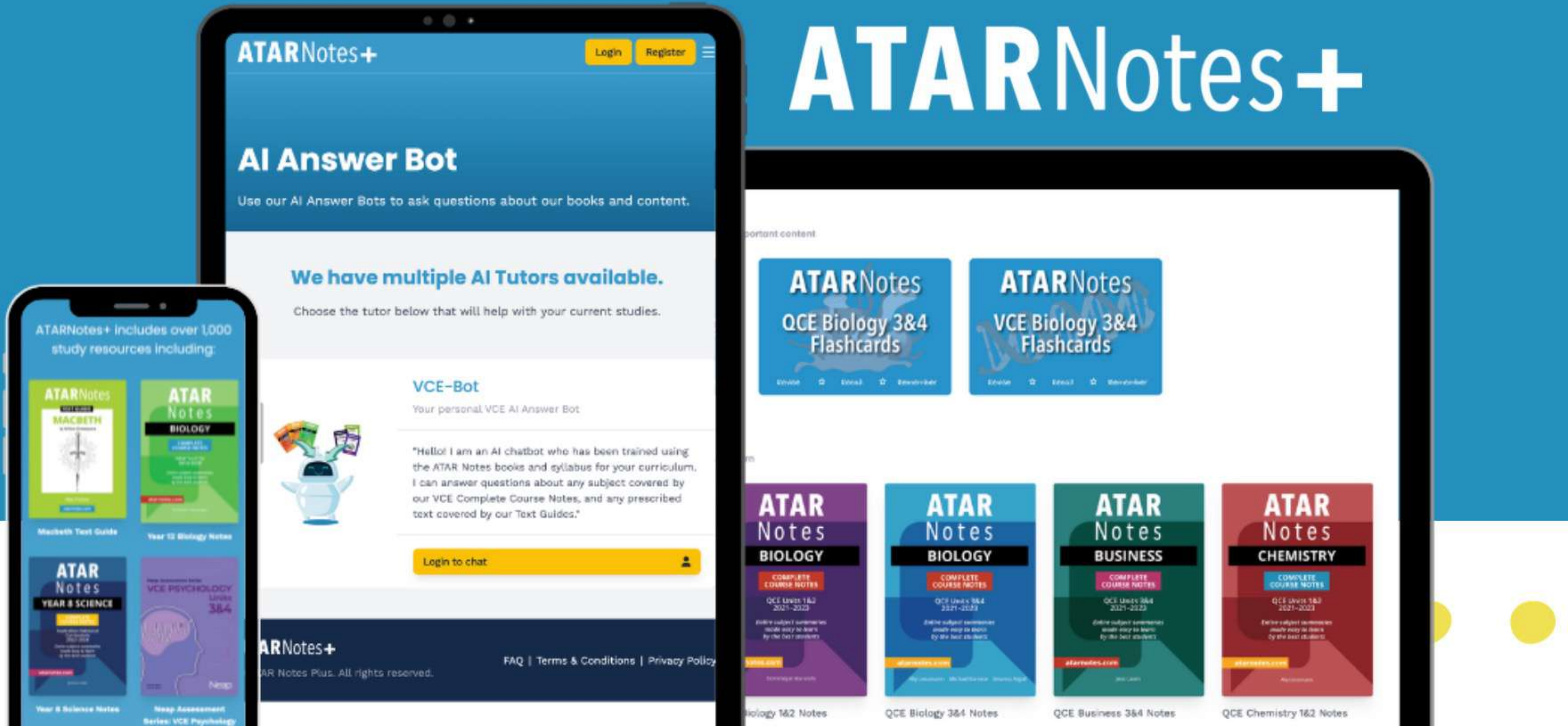


Welcome!!

Topics to be covered

- Welcome to the Chemistry 1/2 head start lecture for 2024 !

House keeping:

- Please feel free to utilise the chat to ask any questions
- The slides should be able to be accessed below
- This recording will be available after the premiere

Who am I?

BLOCK 1:
OVERVIEW & CHEMICAL STRUCTURE

85 minutes

Making a start on the first AOS

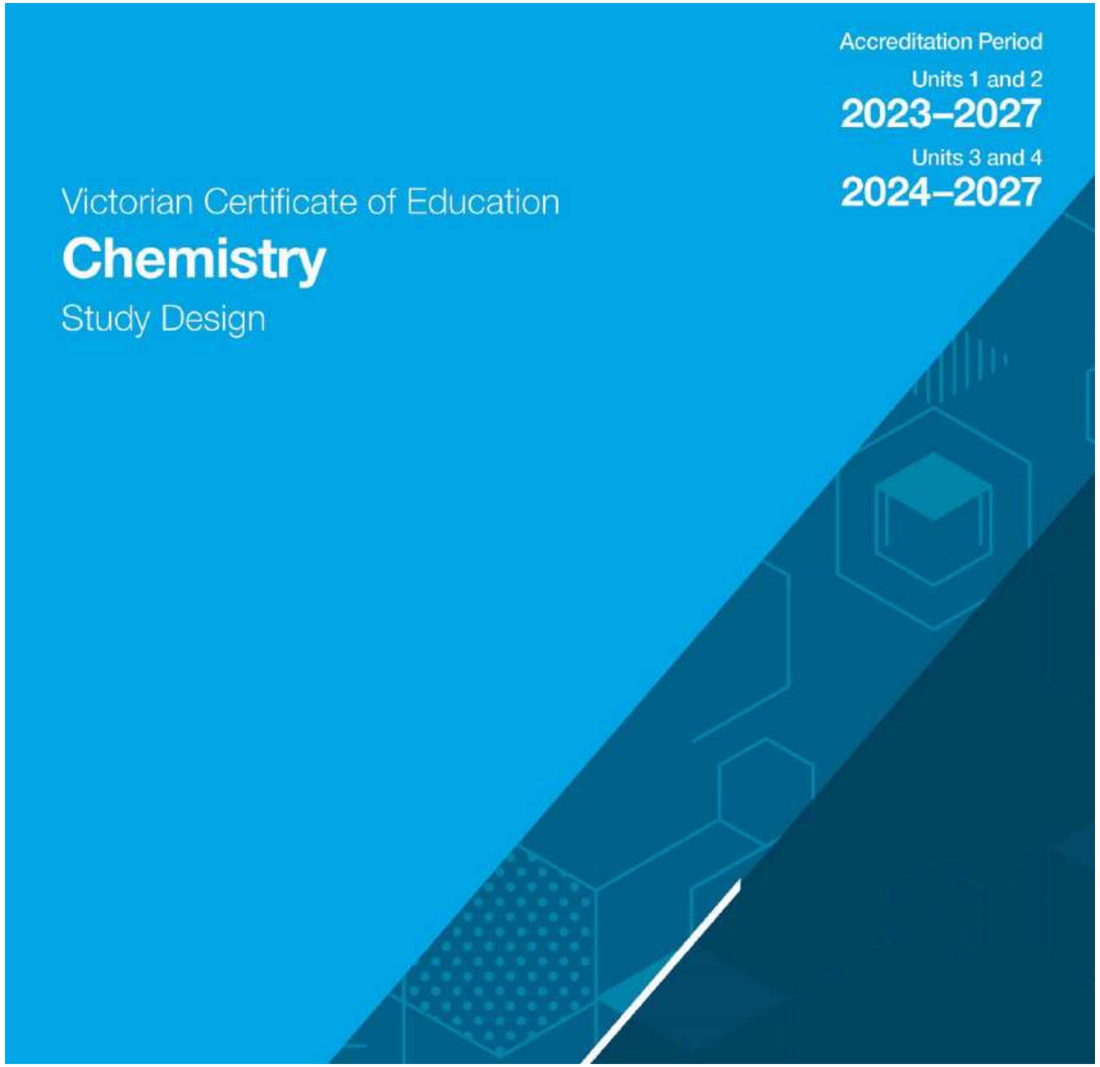
BLOCK 2:
CLASSIFYING CHEMICALS

35 minutes

The start of U1 AOS2 with Calculations and Organic Chem

2023 SD Change

- In 2023 Chemistry converted to a new Study Design
- The 1/2 had quite drastic changes
- Important you utilise the correct SD as the old one is still floating around



VCAA Changes

Added to Unit 1

AOS1:

- Critical elements
- Circular economy of Metal Recycling
- Solubility and precipitations (partially taken from AOS2 of Unit 2)
- Ionic equations (partially taken from AOS2 of Unit 1)

AOS2:

- Relative atomic mass (was already a part of this AOS without being explicitly named in the SD)
- Plant Based Biomass for formation of everyday products
- Polymer Plastics – specifically addition vs condensation reactions with linear vs circular economy
- Fossil Fuel Plastics vs Bioplastics

Essentially, if you come across these questions avoid them

- I have not included content moved to different AOS

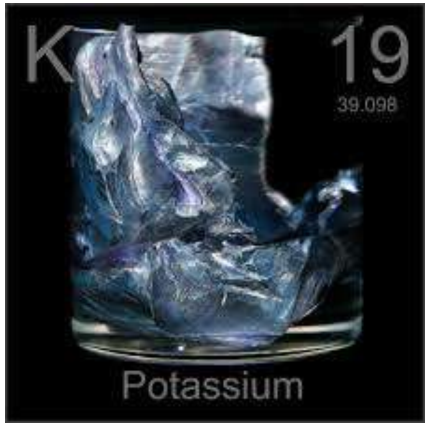
- All nanoparticles and nanomaterials
- S, P, D and F notation of Electron Configuration (Bohr and Schrodinger models)
- Alloy metals
- Origins of Crude Oil

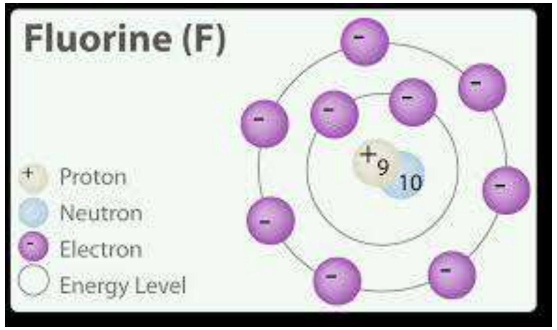
ATAR Notes

Area of Study 1

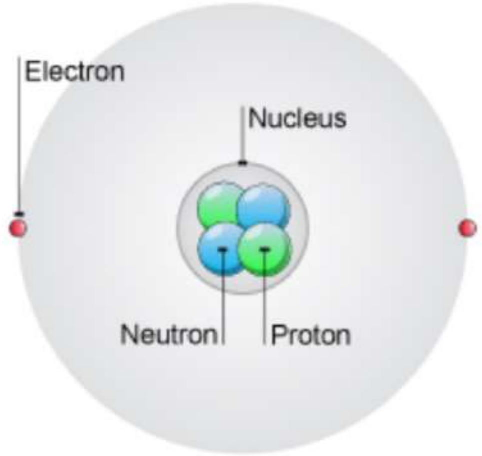
Chemical Structure

- **Atoms** are the basic building blocks of matter
- **Elements** are materials containing one type of atom that cannot be broken down into simpler substances
- **Compounds** are materials containing different types of atoms in fixed ratios
- **Molecules** are substances in which two or more atoms are combined





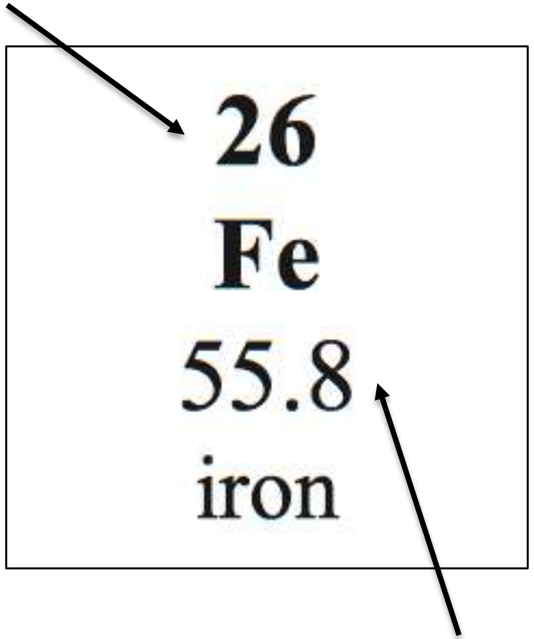
- Atoms are made of subatomic particles, a positive nucleus surrounded by negative electrons
- Nucleus consists of protons and neutrons:
  - Protons are positively charged
  - Neutrons have no charge
- Electrons are negatively charged
  - 1800 times smaller than protons and neutrons
  - Orbit the nucleus in a cloud of electrons
- **Electrostatic attraction** between the nucleus (+) and electrons (–) hold thee atom together



Classifying Atoms

- The number of protons determines the type of atom
- The **atomic number** is the number of protons in the nucleus
- The **mass number** is the number of protons and neutrons in the nucleus
- If the molecule is neutral, the number of protons will equal the number of electrons in the atom

Atomic number = number of protons

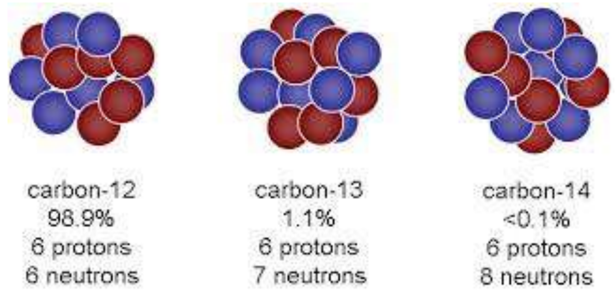


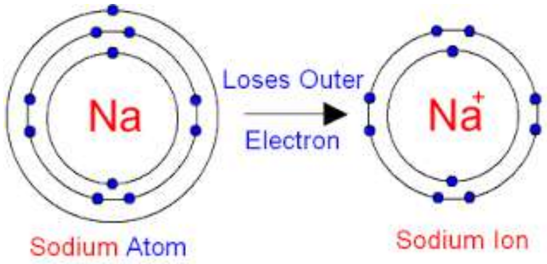
Mass number = number of protons + neutrons

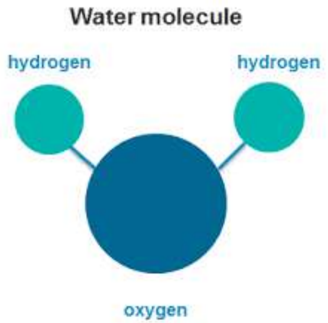
Forms of Atoms

- **Isotopes** are atoms with the same number of protons but different number of neutrons
  - Has altered physical properties

- **Ions** are atoms with a different number of electrons to protons
  - Results in an overall charge
  - Cation (+) and anion (–)







- **Critical Elements** are elements that are thought of as **'endangered'**
  - Think of it just like animals, they are hard to obtain but are extremely desirable in everyday life, there are a few important ones to know

- VCAA specifically names 4 individual / groups of critical elements, they do not expect you to remember specifics but understand a brief summary of each
  - Helium
  - Phosphorus
  - Rare Earth Metals
  - Post Transitional Metals

- Today we will go over the first two briefly

Chemical Structure

Helium

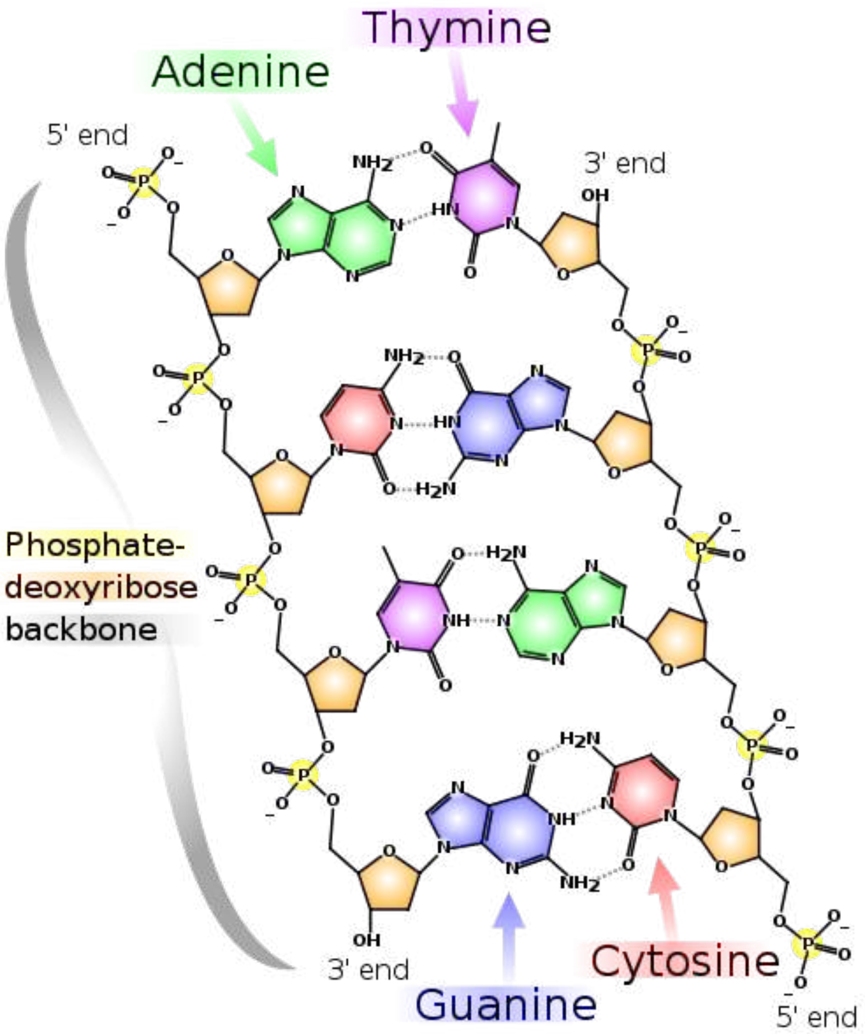
- Number 2 on the periodic table, Helium has an extremely low boiling point and is always obtained as a gas, this make it's useful in various lines of work

- Helium when released as a gas, will rise and rise and rise and rise and rise until it leaves the atmosphere making it a 'truly un-reusable gas'

- Although it is the 2nd most abundant element in our world, it is now known as a critical element

Chemical Structure

- Extremely important in DNA production, phosphorus is vital in cell growth in plants and crops but is used up at considerable rates
- Obtaining phosphorus is even harder, with estimates saying 80% is lost in the chain of production (think 100% obtained initially, only 20% of that will reach its desired use)
- The huge increase in demand on crop growers, has increased the demand for phosphorus rich fertilisers

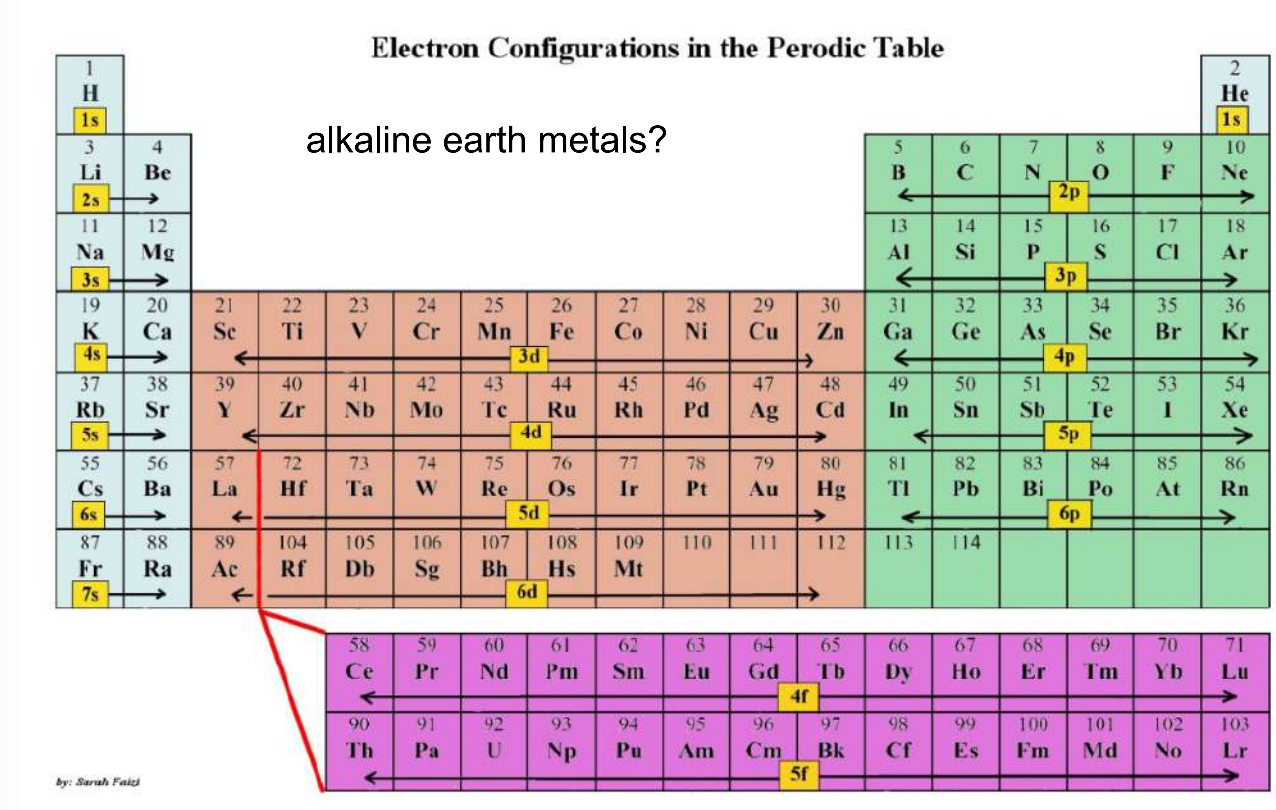


- The periodic table helps us find trends in properties of elements

periods?

halogens?

alkaline earth metals?



alkali metals?

groups?

noble gases?

Modern Periodic Table

- The periodic table is organised in order of increasing atomic number

- Elements are arranged in vertical columns for **groups**
    - Groups determine the number of valance electrons (e.g. Group 1 has 1 valence electron)
    - The number of valence electrons determines many chemical properties, so groups generally have similar properties

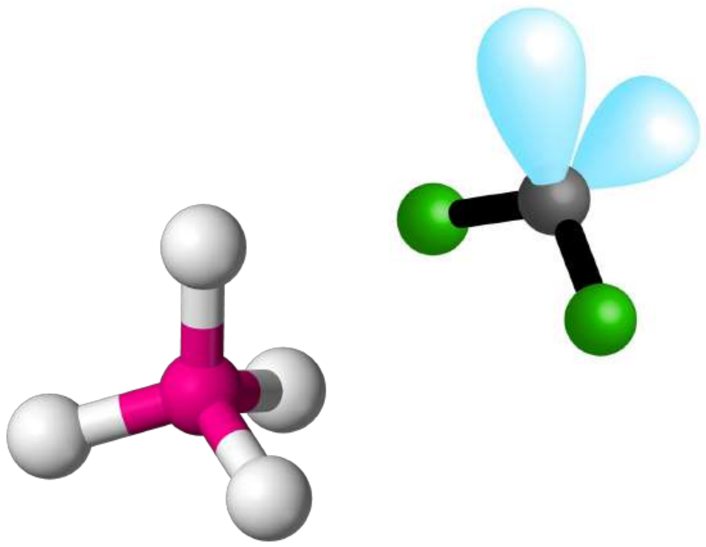
- Horizontal rows are called **periods**
    - The period is equal to the number of occupied shells in the elements atom
    - E.g. Sodium is in Period 3, and it has 3 shells

Key Definitions

- **Core charge** is a measure of the attractive force felt by valence shell electrons to the nucleus
  - $Core\ charge = no.of\ protons\ in\ nucleus\ - no.of\ total\ innershell\ electrons$
- **Electronegativity** is a measure of the ability of an atom to attract electrons towards itself
  - Moving down a group: More shells means further distance between the valance electrons and the nucleus, decreasing electronegativity
  - Moving across a period: As electrons are added to the outer shell, valence electrons are more attracted to the nucleus, increasing electronegativity
- **Atomic radius** is a measurement used for the size of atoms
- **First ionisation energy** is the energy required to remove one valence electron from the outer shell of an element
  - Ionisation is the process of removing an electron from an atom to form an ion – Ionisation forms a cation (positively charged ion)
  - This reflects how strongly valence electrons are attracted to the nucleus

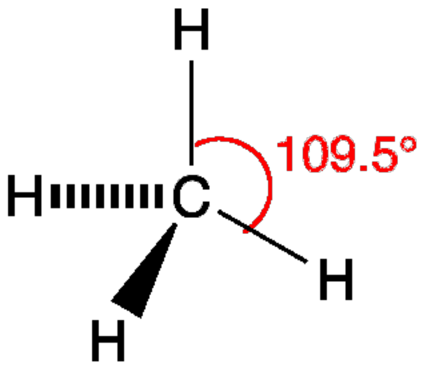
Molecular Shapes

- Different molecules may have different shapes
- Shapes can be predicted using the **valence shell electron pair repulsion theory (VESPR)**
- Like charges repel, so all electrons (which are negative) (lone pairs and covalent bond pairs) will try to be as far away from each other
- Shape is determined by two factors:
  - The number of covalent bonds
  - The number of lone pairs around the central atom

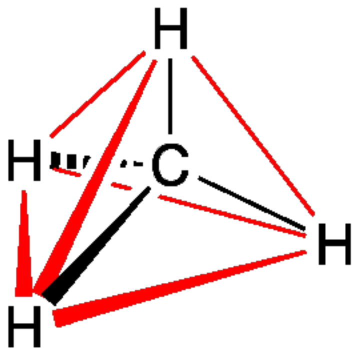


Tetrahedra;

- In cases when there are no lone pairs of electrons and four single bonds, the molecule will form a **tetrahedral** arrangement
- For example, consider methane (CH_4):



The H–C–H bond angle is 109.5°



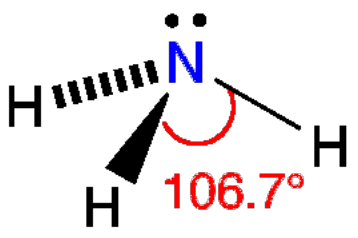
The red lines outline a tetrahedron
Black lines show the covalent bonds

Source: www.chemtube3d.com

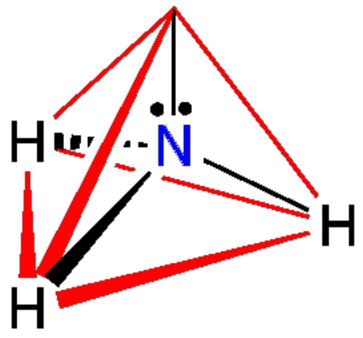
Chemical Structure

Pyramidal

- In cases where there is <u>one lone pair</u> of electrons and <u>three single bonds</u>, the molecule will form a **pyramidical** shape
- For example, consider ammonia (NH_3):



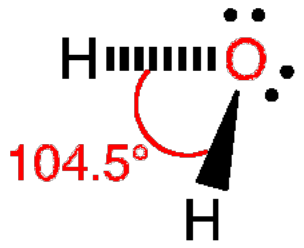
The H–N–H bond angle is 106.7°



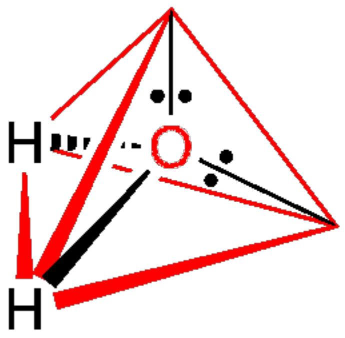
The red lines outline a tetrahedron
Black lines show the electron pairs

Source: www.chemtube3d.com

- In cases where there are two lone pairs of electrons and two single bonds, the molecule will form a V-shape or bent shape
- For example, consider water (H_2O):



The H–O–H bond angle is 104.5°



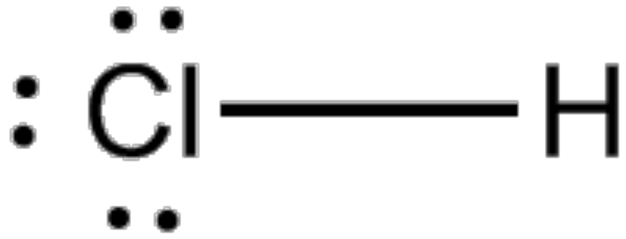
The red lines outline a tetrahedron
Black lines show the electron pairs

Source: www.chemtube3d.com

Chemical Structure

Linear

- In cases where there are three lone pairs of electrons and one single bond, the molecule will form a **linear** shape
- For example, consider hydrogen chloride in gaseous form (HCl(g)):



Bonding and Solubility

Bonding Basics

- Different molecules / compounds have different bonding, indicated by there properties
- Before we look at the type of bonds let's look at some of the common compounds and what their properties say about their bonds

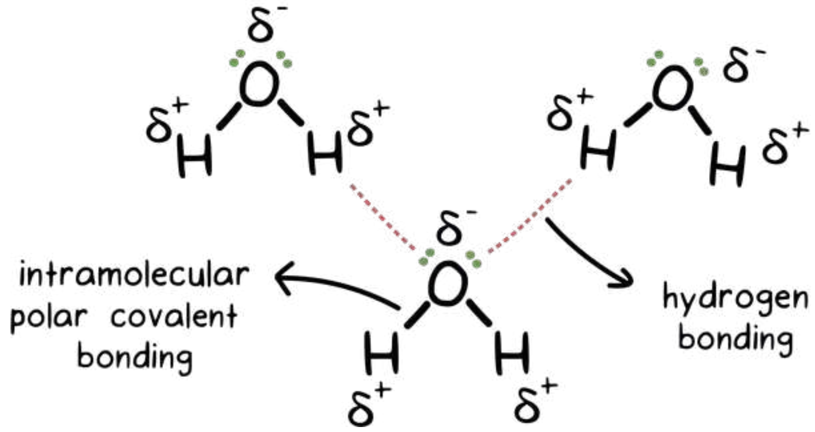
- **Metals** have high melting points and conduct electricity
  - This indicates they have strong bonds; and
  - They contain free-moving charged particles
- **Ionic compounds** have high melting points and conduct electricity only as liquids
  - This indicates they have strong bonds; and
  - They also contain free-moving charged particles when liquid
- **Non-metal compounds** have low melting points and don't conduct electricity
  - This indicates some bonding must be weak; and
  - They do not contain free-moving charged particles

Categorising Bonds

There are two main categories of bonds:

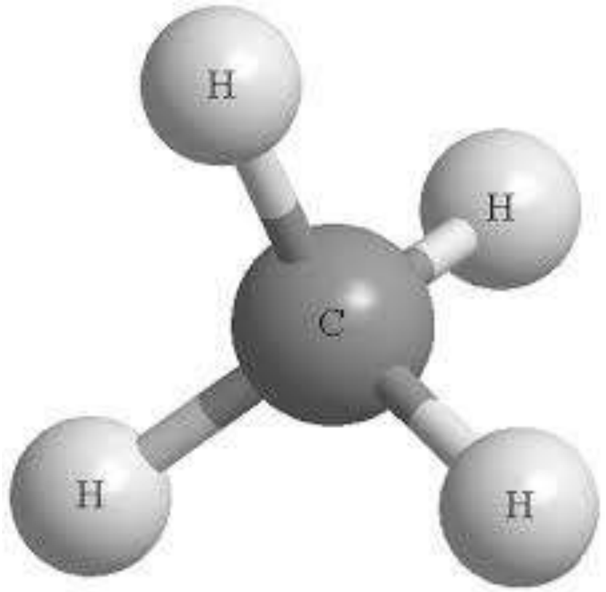
1. **Intramolecular bonds (‘within’)**
    - Forces holding the atoms in a molecule together

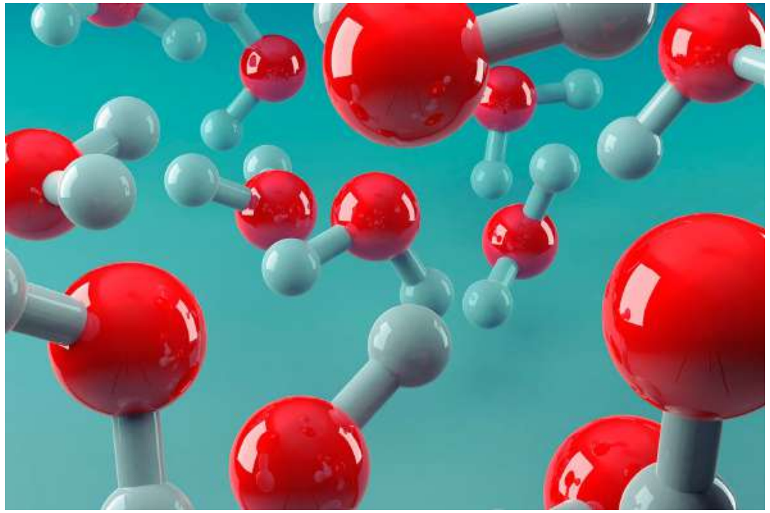
2. **Intermolecular bonds (‘between’)**
    - Forces of attraction holding neighbouring molecules together



INTRA: Covalent Bonds

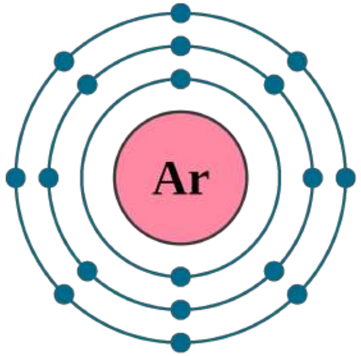
- **Covalent bonds** are a type of intramolecular bond
- The atoms in non-metal compounds are held together with **covalent bonds**, meaning the atoms share electrons to fill their outer shells
- This results in them forming molecules, with different shapes depending on the structure of the outer shell and the atoms involved!

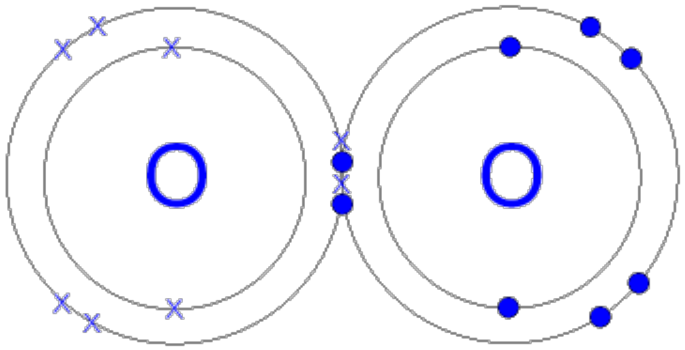




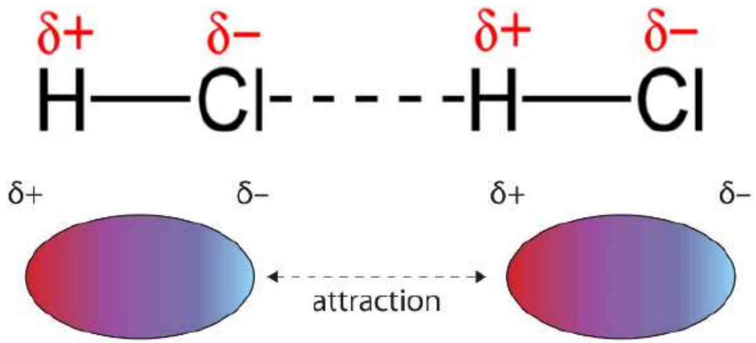
INTRA: Covalent Bonds

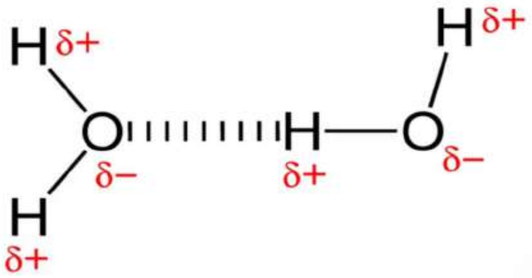
- Non-metals *share electrons*, or *continually exchange electrons* to complete the respective atoms’ outer shells.
- This is known as **covalent bonding**
  - Co = together
  - Valent = valence electrons
- But why do atoms do this in the first place?
  - Noble gases are the most stable of elements due to their full outer shell. As a result, all compounds naturally tend to try and recreate this stability.

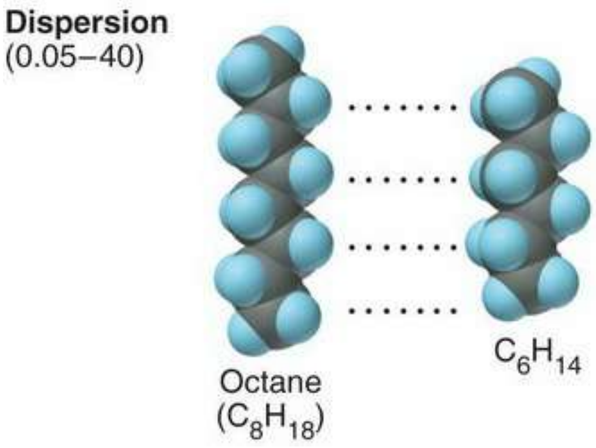




- There are three types of intermolecular forces
  - Dipole-dipole forces
  - Hydrogen bonding
  - Dispersion forces
- These occur between molecules and determine the physical properties of covalent molecular substances

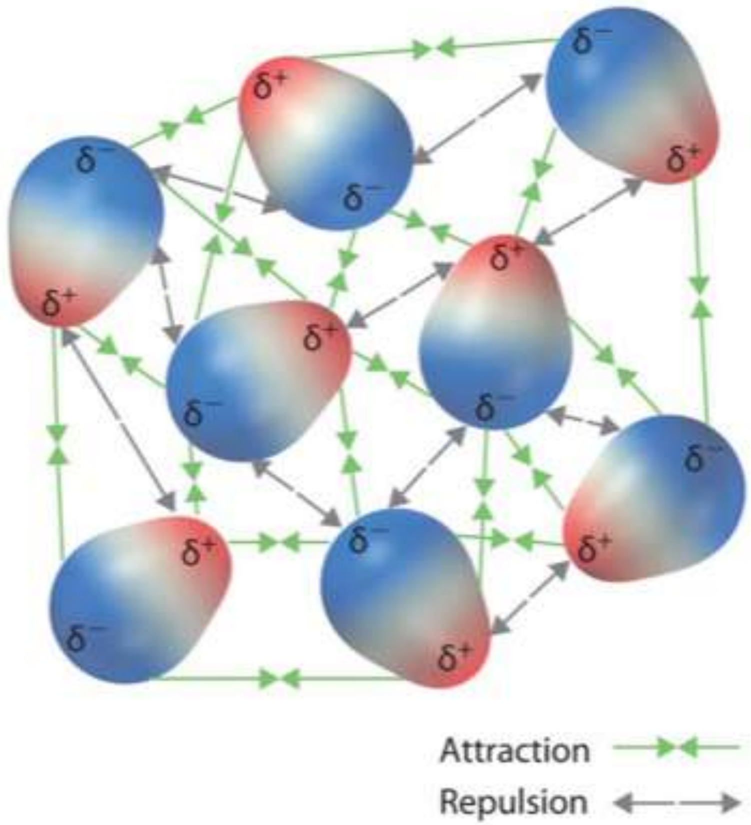






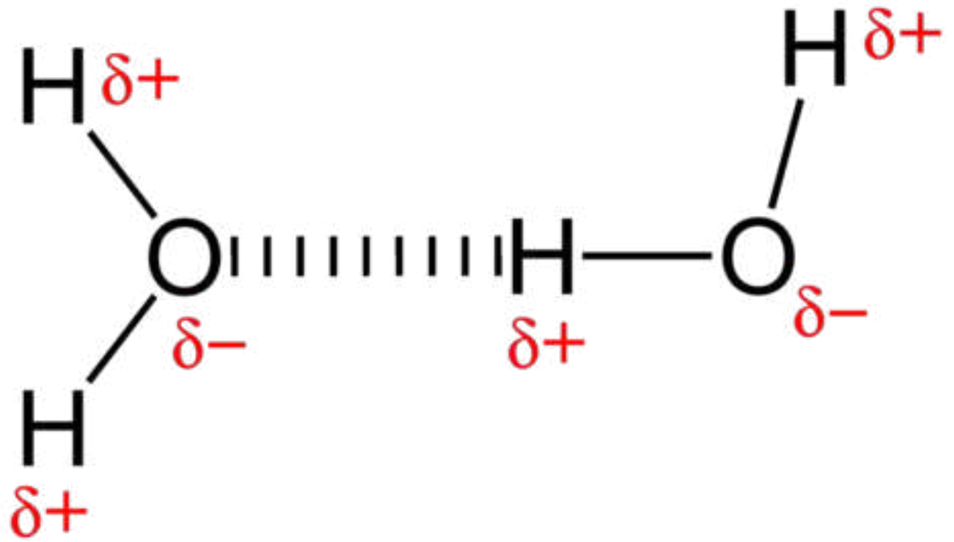
INTER: Dipole-Dipole Forces

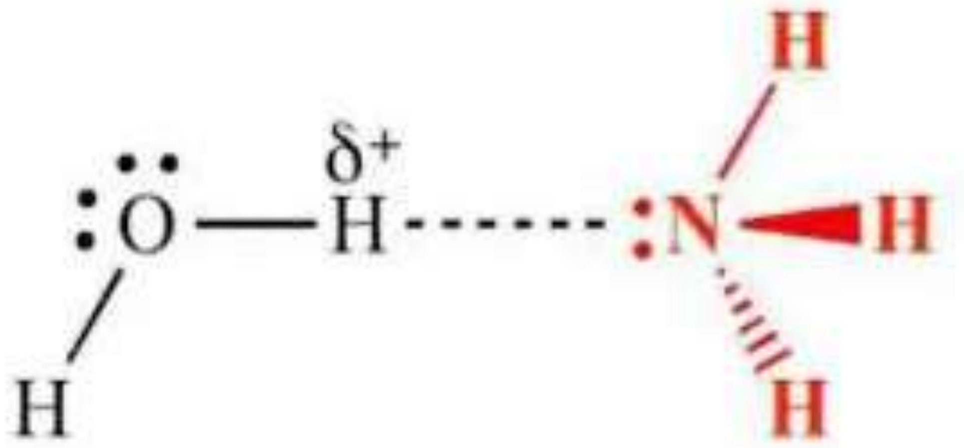
- **Dipole-dipole forces** result from attractions between the partial positive and partial negative sides of **polar molecules**
- They only occur in polar molecules
- They are weak since it is only partial charges (δ– and δ+)
- The bigger the polarity the stronger the attraction is
- Molecules with stronger dipole-dipole forces will have higher melting and boiling points



INTER: Hydrogen Bonds

- **Hydrogen bonding** are a special *type* of dipole-dipole forces that are much stronger
- Hydrogen bonds only occur when a hydrogen bonds with an oxygen, nitrogen or fluorine (**NOF/ FON rule!**)
- Oxygen, nitrogen and fluorine are the <u>smallest</u> and <u>most electronegative</u> elements, so when sharing an electron with hydrogen, they attract the electron the most and create a highly polar bond!





INTER: Dispersion Forces

- There are still forces of attraction between non-polar substances such as oils (which are liquid at room temperature)
- **Dispersion forces** are weak forces of attraction that exist between all molecules (both non-polar and polar)

1. In molecules negative electrons are constantly moving
2. Sometimes they gather more closely to one side
3. This gives that side an **instantaneous dipole**, meaning it gains a **partial negative dipole** for a second
4. This can cause dipoles in neighbouring molecules
5. The attractions between the **instantaneous negative dipoles** and the **instantaneous positive dipoles** are called **dispersion forces**

Properties of Bonding

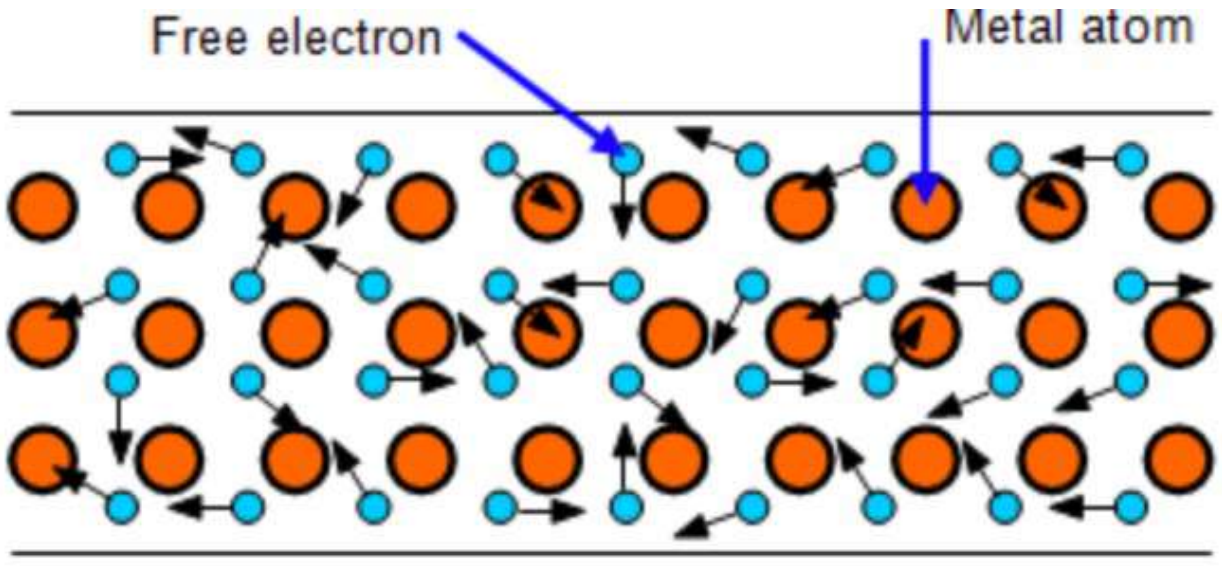
The important properties to know are:

- Boiling point – when increasing the temperature, at what degree do we start to form a gas
  - Water at 100 degrees Celsius begins to boil
- Melting point – when increasing the temperature, at what degree do we start to form a liquid from a solid
  - Water at increasing from 0 degrees Celsius begins to melt
- Strength – how strong the intermolecular bonds between atoms are
- Conductivity – whether electricity can pass through a solid or not, this is determined by whether there are free electrons
- Solubility – ability to dissolve in water

Conductivity

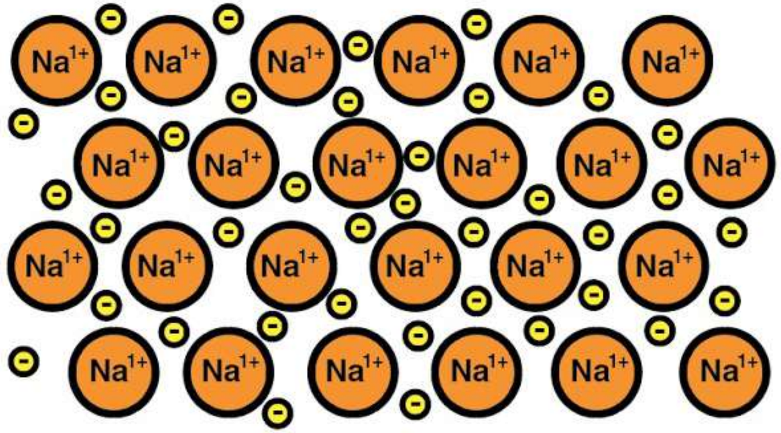
Quick point on conductivity

- It is important to understand that many molecules will be conductive as a liquid, but as a solid it specifically requires free electrons moving within the solid to allow the movement of charge



Conductivity – Metallic Bonds

- **Metals** readily form **cations** (positively charged ions) when they lose electrons to gain a full outer shell
- The freed negatively charged electrons form a ‘***sea of delocalised electrons***’ throughout the metal structure
  - This allows for the conduction of electricity
- **‘Electrostatic force of attraction’** between the positive cations and the negative electrons holds the metal together
- Key point: Use these key terms to explain the properties of metals to gain more marks!



Deducing Structure Based on Properties

- High melting and boiling points = **strong forces between particles**
- Solid at room temperature = **strong forces between particles**
- Are hard but brittle = **strong forces between particles**
- Do not conduct electricity in the solid state = **no free-moving charged particles when solid**
- Good conductors of electricity in the liquid state or when dissolved in water = **free-moving charged particles when liquid or aqueous (Na^+ and Cl^-)**
- Vary from very soluble to insoluble in water. They are not soluble in non-polar solvents such as oil = **mostly polar, though to varying degrees**

Bonding and Solubility

Compounds

- As we discussed earlier, we can understand bonding based on properties and different compounds have different properties

- As a part of AOS1, VCAA wants us to know specifically about ionic and metallic compounds
  - The properties
  - What they made up of
  - How they can be utilised

Bonding and Solubility

Metal Substances

- Where are metals on the periodic table?
- What does their shell structure tell us about their properties?

The Periodic Table of the Elements

1 H Hydrogen 1.00794																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012182											5 B Boron 10.811	6 C Carbon 12.0107	7 N Nitrogen 14.00674	8 O Oxygen 15.9994	9 F Fluorine 18.9984032	10 Ne Neon 20.1797
11 Na Sodium 22.989770	12 Mg Magnesium 24.3050											13 Al Aluminum 26.981538	14 Si Silicon 28.0855	15 P Phosphorus 30.973761	16 S Sulfur 32.066	17 Cl Chlorine 35.4527	18 Ar Argon 39.948
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.955910	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938049	26 Fe Iron 55.845	27 Co Cobalt 58.933200	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.723	32 Ge Germanium 72.61	33 As Arsenic 74.92160	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80
37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.90585	40 Zr Zirconium 91.224	41 Nb Niobium 92.90638	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.90550	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.760	52 Te Tellurium 127.60	53 I Iodine 126.90447	54 Xe Xenon 131.29
55 Cs Cesium 132.90545	56 Ba Barium 137.327	57 La Lanthanum 138.9055	72 Hf Hafnium 178.49	73 Ta Tantalum 180.9479	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.078	79 Au Gold 196.96655	80 Hg Mercury 200.59	81 Tl Thallium 204.3833	82 Pb Lead 207.2	83 Bi Bismuth 208.98038	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (263)	107 Bh Bohrium (262)	108 Hs Hassium (265)	109 Mt Meitnerium (266)	110 (269)	111 (272)	112 (277)	113	114				
				58 Ce Cerium 140.116	59 Pr Praseodymium 140.90765	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.92534	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93032	68 Er Erbium 167.26	69 Tm Thulium 168.93421	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.967
				90 Th Thorium 232.0381	91 Pa Protactinium 231.03588	92 U Uranium 238.0289	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)

Properties of Metal Substances

Metallic Substances:

- **Melting point** = wide range, generally high
  - However, they nearly all have a high boiling point
- **Malleability** = easily manipulated by bending and hammering
  - They are also **ductile**, as they can be stretched into wires
- **Strength** = strong
  - Again, can withstand relatively high amounts of force without being scratched
- **Conductivity** – great conductors of electricity in both solid and liquid form
  - Also, great conductors of heat, think all your household pots and pans
- **Lustre** = the ability to polish to shin is found in metallic substances

Circular Economy

- Circular economy is just a fancy way of thinking about recycling
- If we use copper wiring for an example:
  1. Mine the copper
  2. Extract the pure copper
  3. Form wiring
  4. Utilise the wiring in a product (such as power lines)
  5. Use the power lines until they need replacing / upgrading
  6. Recycle that copper wiring back into pure copper extracts
  7. Utilise that copper to make a new product or even new wiring to then be utilised in a product again

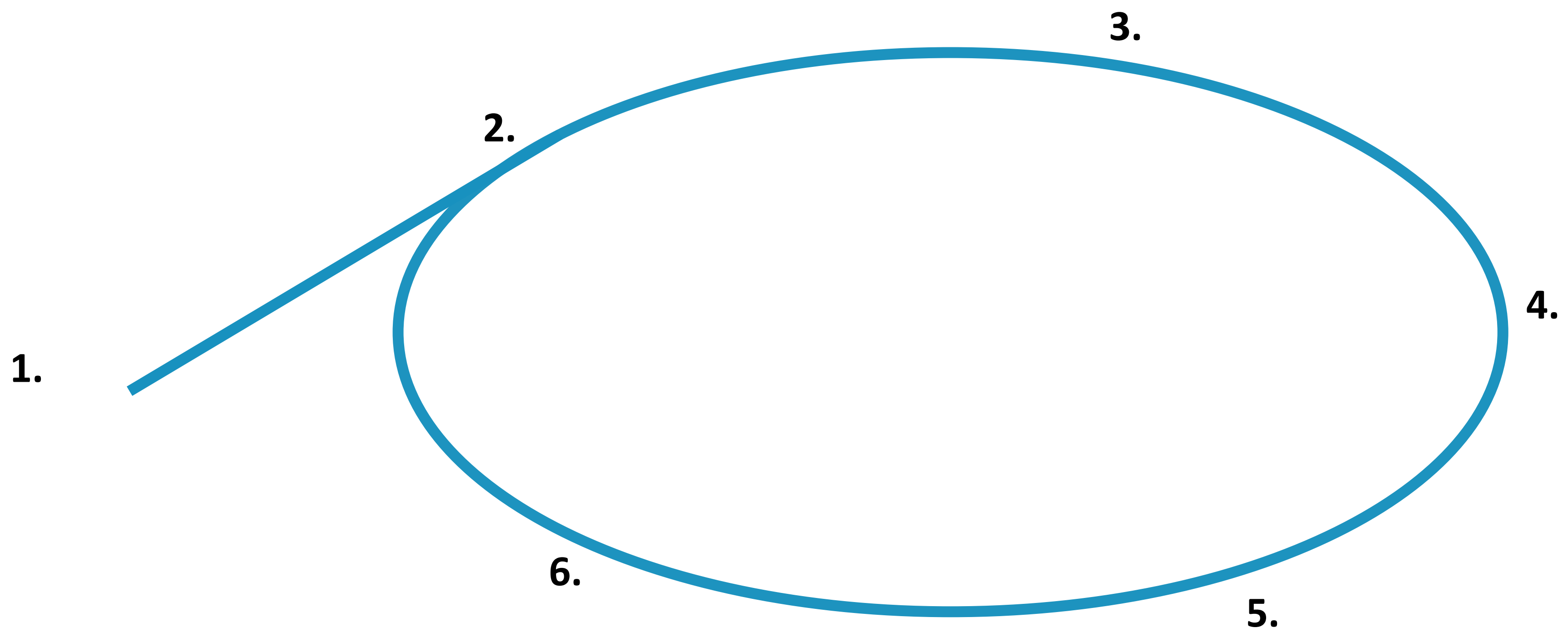
6 Steps of Circular Economy

So, these are the 6 steps you need to utilise whenever answering an exam / SAC question on metal circular economy:

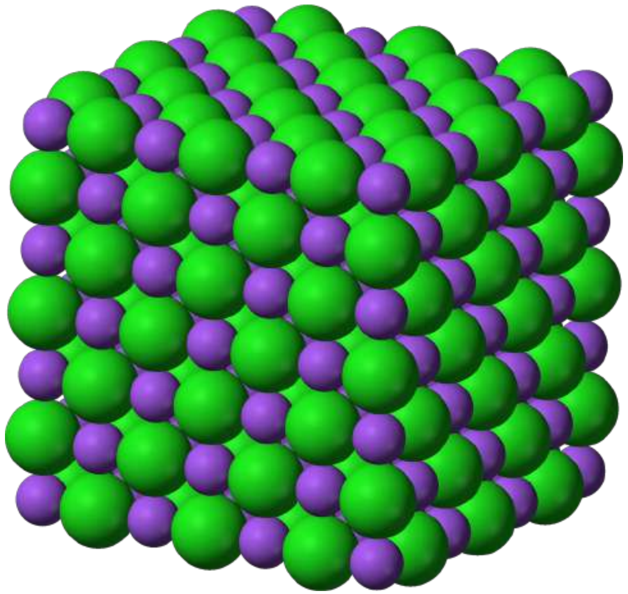
1. Mine the metal
2. Refine the metal
3. Produce a product
4. Utilise the product
5. Recycle the product
6. Reprocess the metal

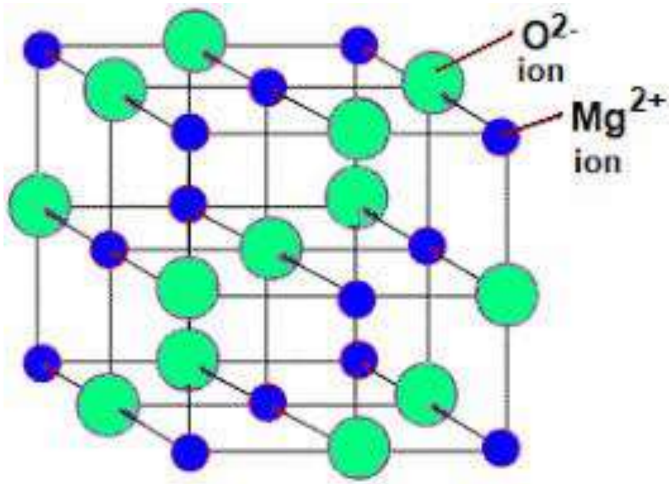
Bonding and Solubility

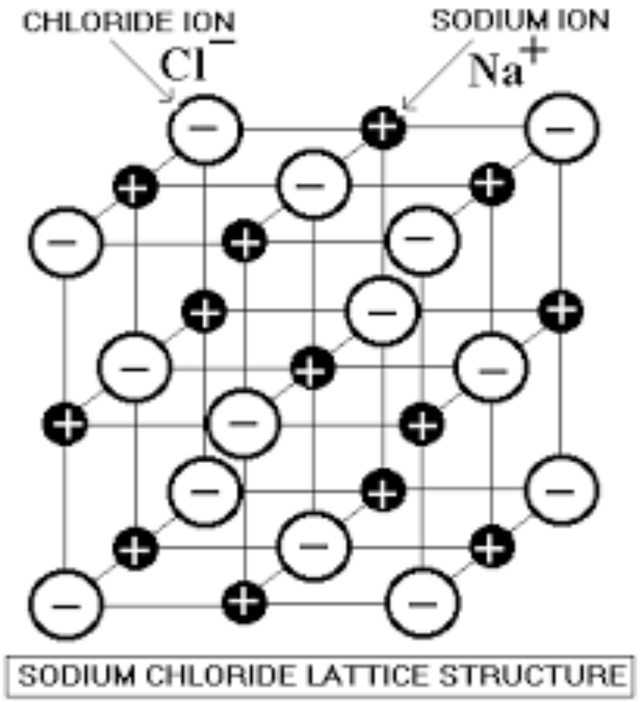
Zinc coating is used on the outside of many ship hulls to ensure the steel structures do not rust. This breakthrough has allowed ships to last longer and for the steel to be easily recycled into new ships. Produce a 6-step cycle for the steel utilized in ships.



- Cations and anions arrange themselves in the following way:
  - Cations and anions combine to form a 3D lattice
  - The 3D lattice is held strongly by electrostatic attraction (ionic bonding)
  - Eg// In NaCl each sodium ion is surrounded by six chloride ions







Properties of Ionic Subtances

Ionic Substances:

- **Melting point** = high
  - Requiring high amounts of energy to be converted from solid state to liquid
- **Malleability** = unable, brittle
  - When struck with a hard strong force, the ionic compound will shatter because layers of ions will move relative to each other
- **Strength** = strong
  - Can be determined by a scratch test, can withstand relatively high amounts of force without being scratched
- **Conductivity** – in liquid state only

Bonding and Solubility

Formulae

1. The charge of most cations can be determined using the periodic table
2. For transition metals, the numerals in brackets indicates the charge

Key anions and cations to memorise

Cations (Positive ions)					
+1		+2		+3	
Hydrogen	H^{+}	Magnesium[1]	Mg^{2+}	Aluminium	Al^{3+}
Lithium	Li^{+}	Calcium	Ca^{2+}	Chromium(III)	Cr^{3+}
Sodium	Na^{+}	Barium	Ba^{2+}	Iron(III)	Fe^{3+}
Potassium	K^{+}	Zinc	Zn^{2+}		
Silver	Ag^{+}	Copper(II)[2]	Cu^{2+}		
Copper(I)	Cu^{+}	Mercury(II)	Hg^{2+}		
Ammonium	NH_4^{+}	Iron(II)	Fe^{2+}		
		Nickel(II)	Ni^{2+}		
		Tin(II)	Sn^{2+}		
		Lead(II)	Pb^{2+}		

Bonding and Solubility

Anions (Negative ions)					
-1		-2		-3	
Hydroxide	OH^-	Oxide	O^{2-}	Nitride	N^{3-}
Hydrogen sulfide	HS^-	Sulfide	S^{2-}	Phosphate	PO_4^{3-}
Hydrogen sulfite	HSO_3^-	Sulfite	SO_3^{2-}		
Hydrogen sulfate	HSO_4^{2-}	Sulfate	SO_4^{2-}		
Hydrogen carbonate	HCO_3^-	Carbonate	CO_3^{2-}		
Dihydrogen phosphate	$H_2PO_4^-$	Hydrogen phosphate	HPO_4^{2-}		
Hydride[3]	H^-	Dichromate	$Cr_2O_7^{2-}$		
Nitrite	NO_2^-				
Nitrate[4]	NO_3^-				
Ethanoate	CH_3COO^-				
Fluoride[5]	F^-				
Chloride	Cl^-				
Bromide	Br^-				
Iodide	I^-				
Permanganate	MnO_4^-				

3. Notice that hydrogen can either be a cation or anion

4. Nitrate is more common in questions than nitrite, but do notice the difference between them.

5. The charge of fluoride, chloride, bromide and iodide can be determined by using the periodic table.

- Our last property of chemicals we look at is solubility and the formation of precipitates in reactions

- Some important definitions for this are:
  - **Dissolution/dissolving**: when you mix a solid into a liquid and the solid disappears into the liquid to become a solution
  - **Solution**: a mixture in which a liquid contains particles that have been mixed into it and disappeared from view
  - **Solute**: what you mixed into the liquid
  - **Solvent**: the liquid in which your solute is being dissolved
  - **Aqueous solution**: a solution where your solute is being dissolved in water

- It's important to understand that polar molecules like polar solvents
- Whereas, yes you guessed it, non-polar molecules like non-polar solvents

- So, if we use water as an example, it is one of the most common solvents and makes solutions aqueous
  - This is due to diluting down the substance
  - Think of it like cordial, you have pure cordial, then add water and it dilutes down the concentration of cordial and therefore, taste of the drink
- Water is a polar molecule and therefore, polar substances will dissolve into the water
- BUT if you add oil, it will just sit as bubbles within the water, as the two do not mix

Dissociate

When ionic compounds dissolve, they ***dissociate***.

- The ionic lattice is attacked by water molecules, where the positive H end attracts the negative anion, and each negative lone pair on the O atom attracts the positive ion.
- This ion-dipole attraction is so strong that the lattice breaks apart, and the ions become surrounded by water molecules (they become hydrated).
- NaCl, NaOH, K_2PO_4 and other compounds all dissolve this way.

$$NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Soluble or Not?

- Not all ionic compounds are soluble though. This is because the energy required to separate their ions from the lattice is greater than the energy released when those ions become hydrated (the solute-solute bonds are stronger than the solute-solvent ones).
- We can use the **SNAPE** rule to see which compounds are *always* soluble.
  - **S**odium (Na)
  - **N**itrates (NO_3^-)
  - **A**mmonium (NH_4^+)
  - **P**otassium (K^+)
  - **E**thanoate (CH_3COO^-)
- Any compound containing any of these 5 ions will always be soluble in water.

- **Precipitation reactions** occur when ions in solution react to form a substance that is **insoluble** in water. That solid substance is called a precipitate. We use solubility rules/solubility tables to work out which is going to form.
- Let's say you have the following two reactants:

$$NaCl(aq) + Pb(NO_3)_2(aq)$$

- If you put these two into the same beaker, the ions will start swimming around. They'll continue doing this until they bump into another ion that they can strongly bond with.
- Thus, you end up having the following reaction take place:

$$2NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$$

Ionic Equations

Now in that last slide we went over precipitates forming, but we it may have been confusing how we knew that ionic equation occurs, so let’s go over the basics of ionic equations

- These generally occur with two salts, a metal with a non-metal, mix and swap their partners, forming two new salts
- It can be easier to go over these using words then the chemical formula

Example:

Sodium Carbonate + Barium Nitrate → Sodium Nitrate + Barium Carbonate

$Na_2CO_3 + Ba(NO_3)_2 \rightarrow 2NaNO_3 + BaCO_3$

Use the following information to answer Questions 2 and 3.

Samples of solid KCl, molten KCl and an aqueous solution of KCl were tested for electrical conductivity.

Question 2

Which of the samples are likely to conduct electricity?

A. molten KCl and an aqueous solution of KCl only
B. solid KCl and molten KCl only
C. solid KCl and an aqueous solution of KCl only
D. all of the samples of KCl

Question 3

Which one of the following statements correctly explains the results of the electrical conductivity experiment?

A. The metal potassium is a component of KCl and all metals conduct electricity.
B. Only the samples of KCl with delocalised electrons will conduct electricity.
C. Any sample that did not conduct electricity must not contain any ions.
D. Charged particles must be able to move freely in order to conduct electricity.

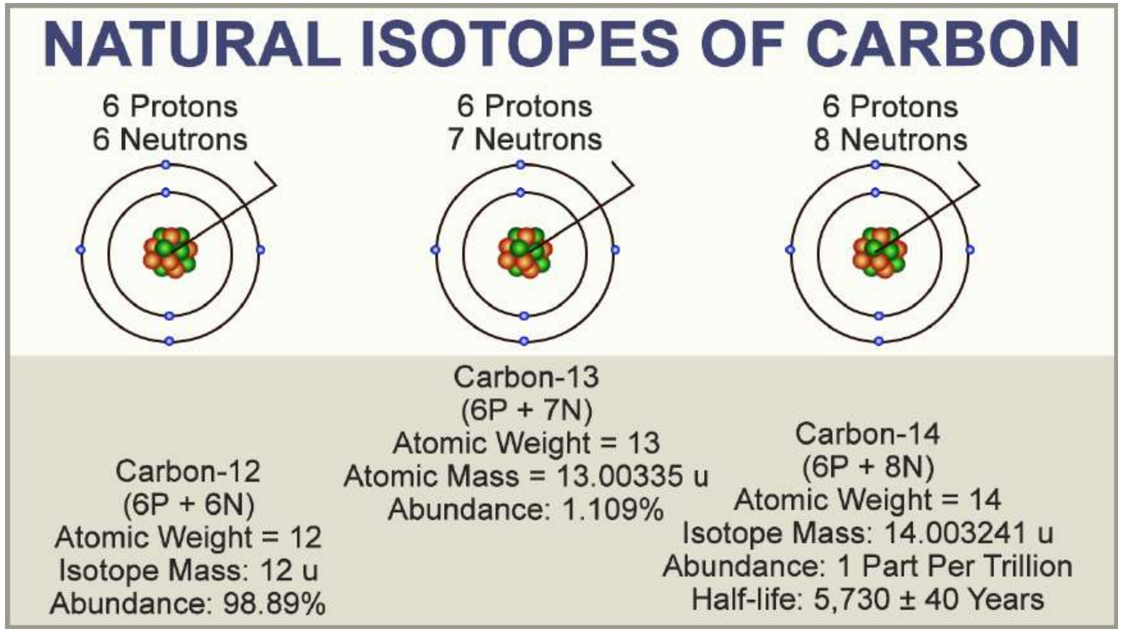
ATAR Notes

Area of Study 2

Classifying Chemicals

Isotopes

- **Isotopes** are atoms of the same element that have different numbers of neutrons (i.e. same atomic number but different mass number)
- The standard to which all masses are compared is the most common isotope of carbon, carbon-12, which is given a mass of 12.



Relative Masses

- Some isotopes are more abundant than others naturally:
- A sample of naturally occurring chlorine consists of:
  - 75% of 35Cl (chlorine with 17 protons and 18 neutrons)
  - 25% of 37Cl (chlorine with 17 protons and 20 neutrons)
- Key point: Electrons aren't considered because they weigh so little

To measure relative isotopic masses and their isotopic abundances we use a **mass spectrometer**

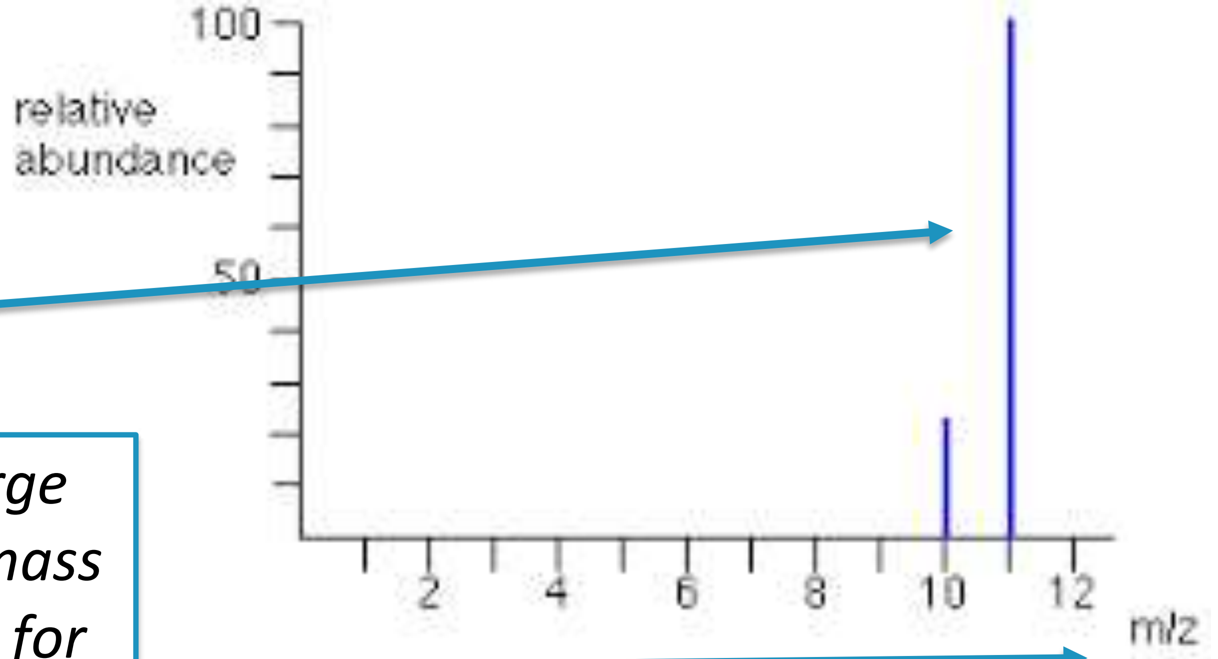
- This gives information regarding:
  - The number of isotopes in a given sample of an element
  - The relative isotopic mass of each isotope
  - The percentage abundance of the isotopes

- The mass spectrometer helps us find out the **percentage abundance** (how much) of each kind of **isotope** (based on the mass number of the given isotope) there are in an element
- It produces a **mass spectrum:**

Relative abundance is the percentage (%) of the given isotope

No. of peaks = no. of isotopes

m/z is the mass per charge ratio - or, the different mass numbers of the isotopes for the given element



Relative Atomic Mass

- All the isotopes of an element cannot be put in the periodic table, so chemists use the relative atomic mass
- The **relative atomic mass** (A_r) is the weighted average of the relative masses of the isotopes of an element on the ^{12}C scale

$$A_r = \frac{(relative\ isotopic\ mass \times \%\ abundance) + (RIM \times \%\ abundance) + \ \ldots}{100}$$

- Where:
  - The number of $(RIM \times \%\ abundances)$ depends on the number of isotopes
  - Note that % abundances must total 100 altogether, so you can use this to calculate the % abundance of an isotope when it is not given to you
  - If the y-axis doesn't give the % abundance, measure each of the peaks and use the following formula:

$$\%\ abundance = \frac{peak\ height}{total\ peak\ height} \times 100$$

PRACTICE QUESTION

Chlorine has two isotopes

- ^{35}Cl with a relative isotopic mass of 35
- ^{37}Cl with a relative isotopic mass of 37

- The relative atomic mass of chlorine is 35.5
- Calculate the percentage abundance of the lighter isotope

Question 6

The formula of a compound is $C_7H_{14}O_2$.

Which one of the following could **not** be the name of the compound?

A. ethyl pentanoate
B. butyl propanoate
C. methyl heptanoate
D. heptanoic acid

The Mole

- Particles in chemicals are so small it would be difficult to count atoms individually
- The quantity for chemists is the mole
- Think of it like how we call 12 eggs a 'dozen eggs'

- Mole is often referred to as 'amount of substance' and is given the symbol ***n*** and the unit ***mol***
  - $n(CO_2) = 2\ mol$ = 'the amount of CO_2 is 2 moles'

- One mole of any substance is the same number of particles as there are atoms in exactly 12 grams of carbon 12
- This is called **Avogadro's number** and is given the symbol N_A
- $N_A = 6.02 \times 10^{23}\ mol^{-1}$
- 1 mol of substance contains 6.02 x 10^{23} particles

So far we have three new quantities:

- The mol (*n*) meaning the 'amount of substance' with the unit mol
- Avogadro's number (N_A) meaning the same number of particles in 12 grams of carbon-12 (6.02×10^{23} particles) with the unit mol^{-1}
- Number of particles (N), meaning the actual number of particles (atoms, ions or molecules)
- The relationship connecting them all together is:

$$n = \frac{N}{N_A}$$

Molar Mass

- Particles of different elements and compounds have different masses
- The **molar mass (M)** is the mass, in grams of one mol of an element or compound, with the unit g mol^{-1}

- One mole is the same number of particles as in 12 g of carbon-12
  - 1 atom of ^{12}C has a relative atomic mass of 12
  - 1 mole of atoms of ^{12}C has a mass of 12 g exactly

- Therefore if:
  - 1 atom of H has a relative isotopic mass of 1
  - 1 mole of H must weigh 1/12th the weight of 1 mole of carbon-12
  - So 1 mol of H weighs 1 gram!

- Another example:
  - 1 molecule of CO_2 has a relative molecular mass of 44
  - 1 mole of CO_2 must weigh 44/12th the weight of 1 mole of carbon-12
  - 1 mole of CO_2 weighs $12 * \frac{44}{12} = 44\ grams$
- From this we can see that the **relative atomic mass** equals the mass of one mol of the atom/elements (or in other words, the **molar mass**)

For example:

- 1 mole of O_2 will equal 32 grams
- This also means that 2 mole of O_2 will equal 64 grams!

- This relationship can be expressed with the following formula:

$$n = \frac{m}{M}$$

- Where,
  - n = amount of mol (mol)
  - m = mass in grams (g)
  - M = molar mass (g mol^{-1})

- Key point: Note there is a difference between capital M and lowercase m
- Now we have two formulas: $n = \frac{N}{N_A}$ AND $n = \frac{m}{M}$
- Key point: Some questions will require you to use both formulas!

PRACTICE QUESTION

1. Calculate the number of CO_2 molecules in 12 grams of CO_2

2. Calculate the number of molecules in 15 grams of O_2

3. Calculate the number of atoms in 4.1 grams of Mg

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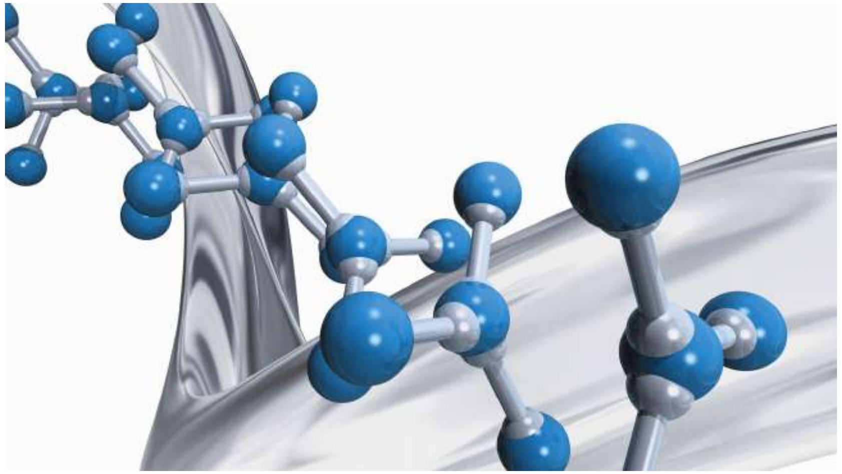
Organic Chemistry

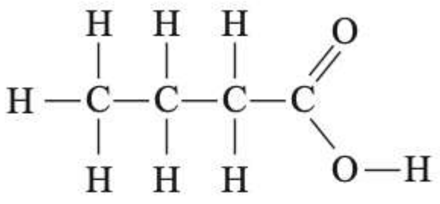
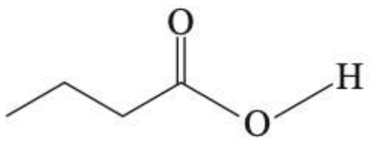
Organic chemistry is the study of compounds of **carbon**

- Carbon is an extremely versatile element:
  - Carbon has four valence electrons that can form covalent bonds with four atoms
  - Carbon atoms can form covalent bonds with other carbon atoms
  - Carbon can form single, double and triple bonds
- All organic compounds contain carbon, but can also contain hydrogen, oxygen, sulphur or chlorine/halogens etc



- Everything else is classified as inorganic
- Hydrocarbons are compounds formed between carbon and hydrogen ONLY
- Hydrocarbons are found in crude oil and are commonly used as fuels

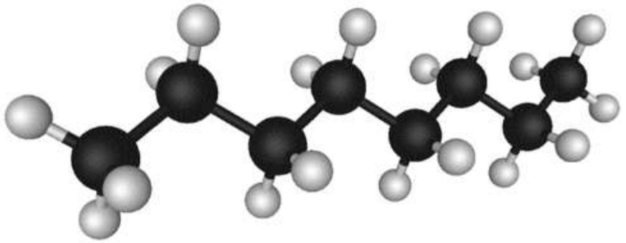
- There are a number of ways to write out formulas:

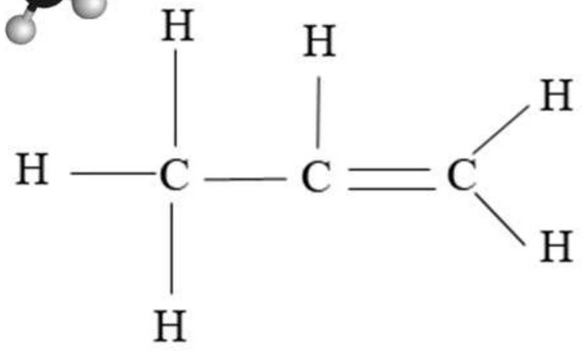
Formula	Representation
molecular formula	$C_4H_8O_2$
structural formula	
semi-structural (condensed) formula	$CH_3CH_2CH_2COOH$ or $CH_3(CH_2)_2COOH$
skeletal structure	

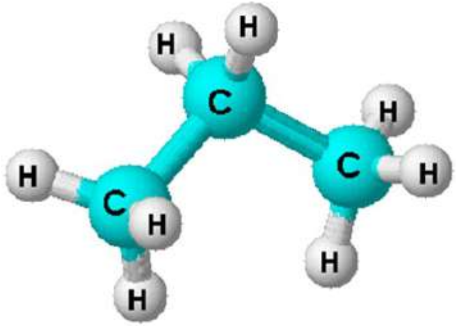
Organic Chemistry

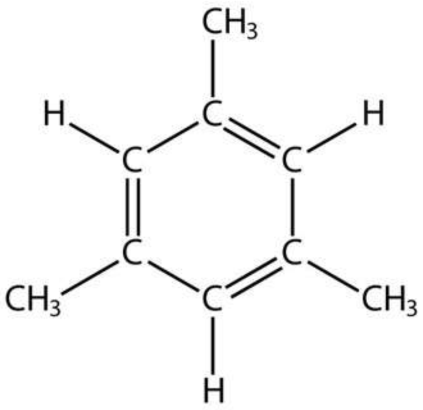
Alkanes

- Hydrocarbons are composed of only carbon and hydrogen
- **Alkanes** are hydrocarbons that contain only single carbon-carbon bonds (e.g. methane CH_4, ethane C_2H_6)
- These are known as **saturated**, meaning they only contain **single carbon-carbon bonds**
- Which of the following compounds are saturated?



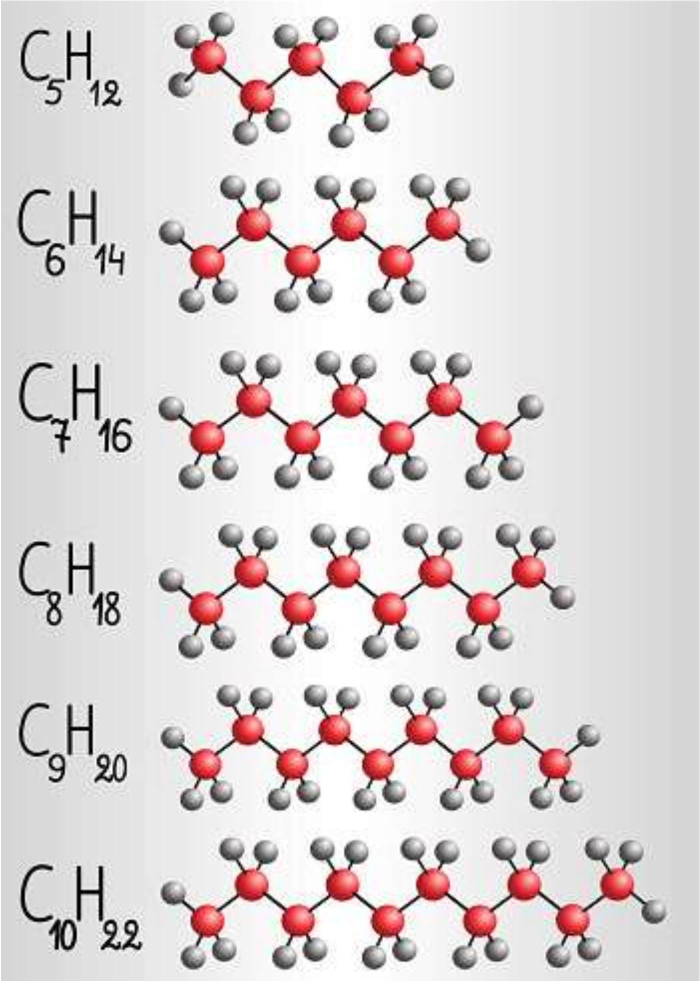






Alkanes

- Each alkane differs to a previous member by $-CH_2-$
- A series of molecules that differ by a CH2 from a previous member is known as a **homologous series**
- Compounds in the same homologous series have:
  - A similar structure
  - A pattern in physical properties
  - Similar chemical properties
  - The same general formula

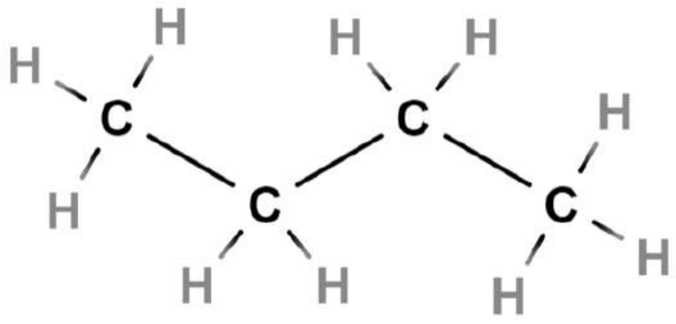


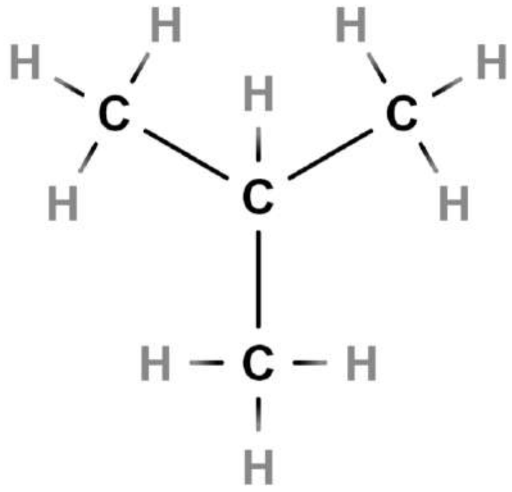
Alkanes share the same general properties:

1. They are non-polar, so are insoluble in water
2. The boiling point of alkanes increases with chain length, as dispersion forces increase
3. They have the general formula: C_nH_{2n+2}

- Alkanes are named by finding the stem name according to the number of carbons (meth–, eth–, prop–, but–, etc) and adding '–ane' after the stem name
- E.g. CH_4 has 1 carbon, so its stem name is meth–, and it is an alkane so its suffix is –ane, so it is called **methane**!

- Molecules that have the same molecular formula but different molecular structures are known as **structural isomers**
- Methane, ethane and propane can only form one structure
- However, C_4H_{10} can form two structures:





- Which of these has a greater boiling point?

Naming Alkanes

Nomenclature – refers to the method in which we name organic chemicals and is used throughout Unit 1, 2, 3 and 4

- Alkanes are named by finding the stem name according to the number of carbons (meth–, eth–, prop–, but–, etc) and adding '–ane' after the stem name
- Eg// CH_4 has 1 carbon, so its stem name is meth–, and it is an alkane so its suffix is –ane, so it is called **methane**!

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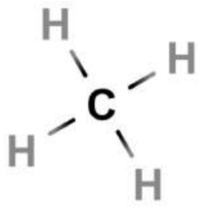
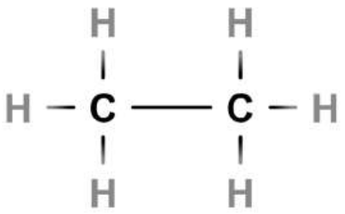
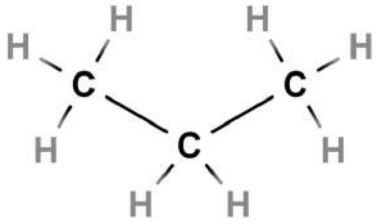
Naming Alkanes

Name	Formula	Structure
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_8	

Stem (parent) name	Number of carbon atoms
Meth–	1
Eth–	2
Prop–	3
But–	4
Pent–	5
Hex–	6
Hept–	7
Oct–	8
Non–	9
Dec–	10

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Under the IUPAC system of naming, the following rules apply when naming alkanes:

1. Identify the longest unbranched carbon chain
2. Number the carbon atoms in the chain from the end of the chain that will give the smallest number to branching groups
3. Name the alkyl group after the alkane from which they are derived
4. Place the number and position of each alkyl group at the beginning of the compound's name
5. If there identical side chains are present, use 'di-' as a prefix; for three use 'tri-'.
6. If there are alkyl side chains of different lengths on the molecule, list them in alphabetical order at the start of the name, with their numbers to indicate their respective positions

ATAR Notes

GOOD LUCK <3